Periodic table of elements

History of the periodic table of elements

- First attempt to divide the elements - German chemist **Johann Döbereiner** - 19th century - triad rule - in a triad of elements (e.g. Li, Na, K) the middle element has the average properties of the outermost elements.
- 2nd half of the 19th century - English chemist **John Newlands** - rule of octaves - first arrangement of elements according to their atomic weight (he used the analogy of the similarity of elements with the similarity of tones in music).
- **Dmitry Ivanovich Mendeleev** published his Periodic Table in the Journal of the Russian Chemical Society in 1869.
  - He ranked the elements according to their relative atomic weight (63 elements were known at the time). He formed groups of elements with similar properties, breaking the list into a series of rows and columns, thus creating the periodic table of elements.
  - The original definition of the Periodic Law: **The properties of the elements are a periodic function of their atomic weights.**
  - Mendeleev left spaces in the table for elements yet to be discovered and also predicted their chemical properties with high accuracy.
  - The gradual discovery of further knowledge proved that the Periodic Law is *one of the fundamental laws of nature.*
  - Today, based on the knowledge of the internal structure of atoms, the periodic law is formulated as follows: **The properties of the elements are a periodic function of their proton numbers.**

Periodic table of elements

- The short and long forms (the most commonly used, it also contains lanthanoids and actinoids), of the non-traditional modifications, the main example is the table of elements according to their occurrence on Earth.
- The elements in the periodic table are arranged in **7 horizontal series - periods**, which are denoted by Arabic numerals, and in **18 vertical groups** denoted by Roman numerals. The groups are further divided into **major subgroups** (I.A...VIII.A) and **minor groups** (I.B...VIII.B).
- **The period number is the same as the maximum principal quantum number**, i.e. the number of the valence layer of electrons.
- The first period begins with hydrogen H, which has an electron configuration of 1s1, and ends with helium, which has an electron configuration of 1s2
- The element of each subsequent period begins with the electron configuration ns1
- According to the filling of certain types of orbitals with valence electrons, the elements can be divided into:
  - **non-transitional elements:**
    - **s-elements** - valence electrons are in ns orbitals
    - **p-elements** – valence electrons are in ns np orbitals
  - **transitional elements**:
    - **d-elements** – valence electrons are in ns (n – 1)d
  - **internally transient elements**
    - **f-elements** - lanthanoids and actinoids - valence electrons complement the (n - 2)f orbitals
- **the number of valence electrons is identical to the number of the group** in which the element is located (exception - some elements of group VIII.B)
- From the position of the element in the table we can determine:
  - the electron configuration of the atom
  - the physical and chemical properties of the element (they are periodic in nature)
  - many properties and possible compounds can be predicted

Periodicity of physical and chemical properties of atoms of elements

First ionization energy of atoms of elements

- the energy required to remove one electron from each atom in 1 mole of gaseous atoms in the ground state
- **the ionisation energy value characterises the ability of an atom to give up electrons**
- the main factors affecting the 1st ionization energy:
  - **the size of the positive charge on the nucleus of an atom** - the increase in the nucleus of an atom (with increasing proton number) will also cause an increase in the 1st ionization energy - in layman's terms: the larger the nucleus, the larger the ionization energy
  - **the distance of the electron from the nucleus** - **the further** the valence layer of electrons is **from the nucleus**, **the less ionization energy** is required to detach the electron
  - **electron capping effect** - refers to atoms that contain inner electron layers. The inner electrons "dampen" the effect of the attractive nucleus. **The greater the covering effect**, **the less ionization energy** is needed to strip the electron
- Hence: **the first ionization energy of atoms of elements increases from left to right in periods and decreases downwards in groups**
- while the combined effect of distance and overlap weakens the effect of increasing the positive charge of the nucleus

Atomic radii of elements

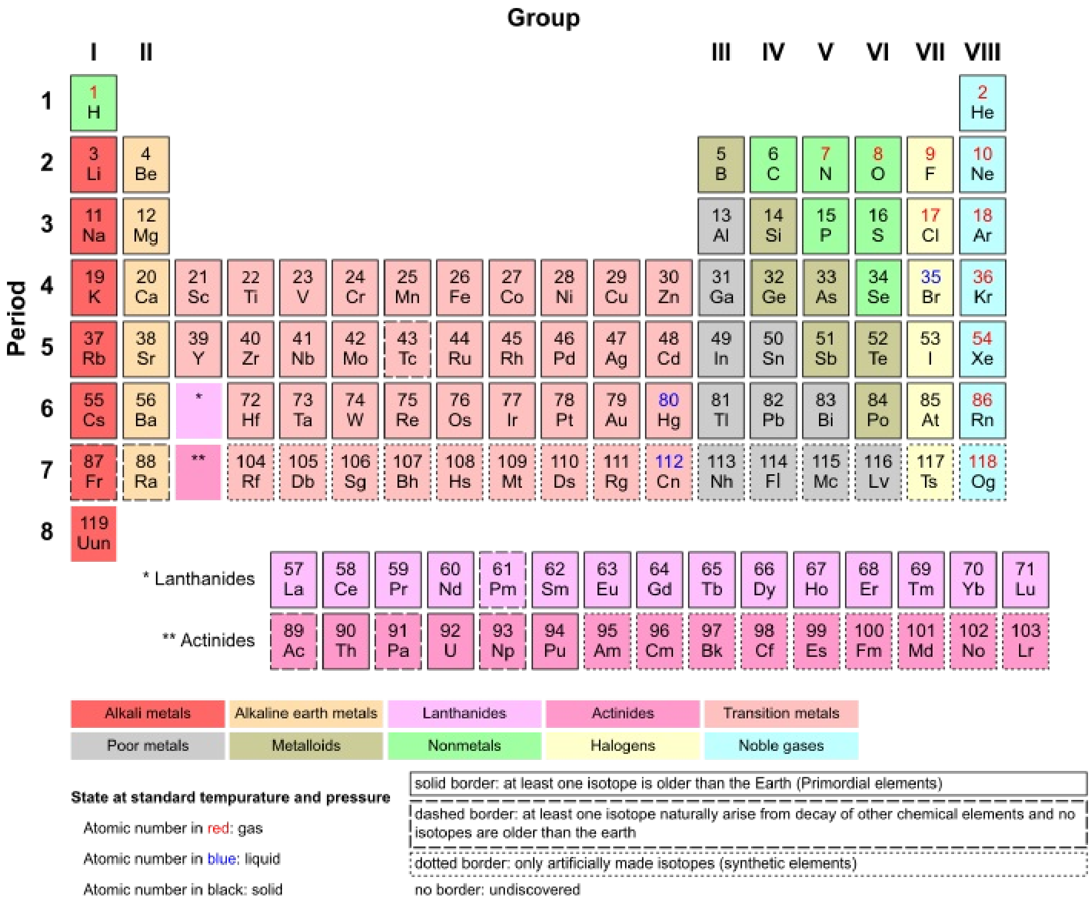
- atomic radii of the elements in the period decrease from left to right and in groups increase downwards

Elektronegativity

- Electronegativity is the ability of an element to attract binding electrons to itself (a measure of its ability to attract its binding electrons)
- **it rises from left to right in a period and falls downward in a group**
- electronegative elements: trying to reach the electron configuration of the nearest noble gas - the so-called **octet rule**
- electropositive elements: trying to reach the configuration of the preceding noble gas

Metals, non-metals and semi-metals

- **metals** - electropositive elements, mainly s-elements and p-elements with a small number of electrons in the valence layer, d and f-elements. They easily form cations, in the solid state they form a metallic lattice
- **non-metals** - electronegative elements, especially p-elements with a larger number of electrons in the valence layer (for example: halogens, chalcogens, hydrogen, carbon, nitrogen,...). They easily form anions.
- **Semi-metals** - have some properties of metals and some of non-metals.
- **Metallic properties of elements increase in groups downwards and in periods from right to left**



References

Related articles

- Atom
- Atomic nucleus

Literature used

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